ANSWERS: Explaining Equilibrium

1) When a change is made to a system that is at equilibrium, the system responds to reduce the effect of that change. If there is an increase in pressure, the system responds by decreasing the pressure. This occurs by favouring the reaction that produces fewer gas particles. Because there are now fewer particles hitting the sides of the container, there is less pressure.

In **Reaction One** there are two moles of gas particles on each side of the equation. Because there are the same numbers of gas particles on both sides of the reaction, then a change in pressure will have no effect as neither reaction will be favoured. In **Reaction Two** however, there are four moles of gas particles on the reactant side of the equation and two moles of gas particles on the product side of the equation. Therefore, when there is an increase in pressure, the system would shift and favour the forward reaction meaning there are now fewer gas particles overall and hence fewer gas particles hitting the sides of the container and therefore less pressure overall.

2) When $PCl_3(g)$ is removed:
Amount of Cl_2 increases.
As $PCl_3(g)$ is removed / concentration decreased, the equilibrium will shift to oppose the change, i.e. increase the concentration of $PCl_3(g)$.
This will favour the forward reaction, producing more Cl_2.

When the pressure is decreased:
Amount of Cl_2 increases.
Decrease in pressure causes the equilibrium to shift to increase the number of gaseous particles, i.e. shifts equilibrium to the side with the greatest number of moles. Since there are two moles of gaseous products and one mole of gaseous reactant, equilibrium will shift to right.
This will favour the forward reaction, producing more Cl_2.

3) At increased temperature the value of K_c increases.
This means that equilibrium shifts in favour of products i.e. the forward direction.
An increase in temperature causes the equilibrium to shift to favour the reaction that absorbs heat / energy, i.e. the endothermic direction.
Hence, the forward reaction is endothermic.

4 i) The colour of the solution becomes less brown.

Decreasing the temperature causes an equilibrium shift to favour the reaction that releases energy / heat, ie shift in the exothermic direction. This is the reverse direction (favouring N_2O_4).

ii) The colour of the solution becomes more brown.

Decreasing the pressure causes the equilibrium to shift to increase the number of gas particles in the equilibrium mixture, ie shift to the side with the greater number of moles. This is the forward direction (favouring NO_2).

5) $Kc = \frac{[CH_3OH]}{[CO][H_2]}$ *Kc* small OR *Kc* <1: reactants > products

When temperature is decreased: reaction occurs to: reduce decrease in T: favouring the forward, exothermic reaction.
A catalyst has no effect on yield: since it increases the rates of forward and reverse reactions equally.
When pressure is increased: reaction occurs to: reduce increase in pressure: favouring the forward reaction which reduces the number of moles of gas.

6) HI(*g*) added.
(i) Colour will become more purple.
(ii) As the concentration of HI is increased, the equilibrium will shift to oppose the change, ie decrease the concentration of HI.
This will favour the reverse reaction producing more I_2, so more purple.
Mixture is cooled.
(i) The purple colour fades.
(ii) Decreased temperature causes the equilibrium to shift to favour the reaction that releases energy / heat, to replace the heat that has been lost. ie the exothermic direction.
This will favour the forward reaction resulting in less I_2 so less purple.
Increase in pressure.
(i) No change.
(ii) Increase in pressure causes the equilibrium to shift to reduce the number of **gaseous** particles, ie shifts equilibrium to the side with the least number of moles. Since each side of the equilibrium equation has two moles there will be no equilibrium shift. The colour will remain the same.

7) a i) The forward reaction is exothermic. A decrease in temperature causes an equilibrium shift to favour reaction that releases energy, ie shift in the exothermic direction. So to have a greater amount of $SO_3(g)$ in the equilibrium mixture, the temperature must be low.

ii) The lower the temperature used, the slower the reaction rate.

Although a greater amount of $SO_3(g)$ will be present in the equilibrium mixture, it will be uneconomical if it takes a long time for the reaction to reach that equilibrium.

Approximately 450°C is a compromise temperature producing a sufficiently high proportion of sulfur trioxide in the equilibrium mixture, but in a short time.
b i) Use a high pressure / decreasing volume.

ii) High pressure / decreasing volume

There are 3 **gaseous** moles / molecules on the left-hand side of the equation, but only 2 moles / molecules on the right.

If the pressure is increased, the system will move to minimise the effect of this and favour the reaction that produces fewer molecules of **gas**, since that will cause the pressure to fall again, ie, will favour formation of $SO_3(g)$.

8) Fizziness: As the lid is opened, $CO_2(g)$ escapes from the drink and the **pressure** is decreased. The equilibrium in **Equation One** will shift to the left. The position of equilibrium moves to minimise the effect of the change.
ie, the decrease in **pressure** favours formation of more moles / molecules of gas, so the position of equilibrium will move to favour formation of more $CO_2(g)$ in Equation One.
This results in a lower concentration of $CO_2(aq)$. As more $CO_2(aq)$ is lost from the drink, there is less fizz in the drink.
pH: As the concentration of $CO_2(aq)$ is decreased, the position of equilibrium in **Equation Two** will shift to favour formation of reactants, ie form more $CO_2(aq)$.
As this occurs, the concentration of H_3O^+ (and HCO_3^-) ions decreases. As the $[H_3O^+]$ decreases, the pH will increase.

9) a) Colour of solution goes orange / becomes a lighter orange / lighter red / colourless
Decreased temperature causes an equilibrium shift to favour reaction that releases energy / heat, ie shift in the exothermic direction.
As forward reaction is endothermic having a positive $\Delta_r H$, the reverse reaction is exothermic.
Equilibrium shifts in exothermic / reverse direction and the concentration of $FeSCN^{2+}$ will be decreased, so colour of solution is lighter.

b) Colour of solution goes orange / becomes a lighter orange /lighter red / colourless

As the concentration of Fe^{3+} ions is decreased (because of reaction with fluoride ions) equilibrium will move to increase the concentration of Fe^{3+} ions (a reactant).

So reverse reaction is favoured, the concentration of $FeSCN^{2+}$ will be decreased so that colour of solution is lighter.

c) Colour of solution goes dark red / becomes a darker red.

As the concentration of Fe^{3+} ions is increased (due to $FeCl_3$ dissolving), equilibrium will move to decrease the concentration of Fe^{3+} ions (a reactant).

So forward reaction is favoured, the concentration of $FeSCN^{2+}$ will be increased so that colour of solution is darker.

10) $K_c = \dfrac{[NH_3]^2}{[H_2]^3[N_2]}$

Line A

- Equilibrium shifts to reduce pressure increase.
- Product side has fewer moles.
- Equilibrium shifts in forward dir./least amt dir.

Therefore increase in % NH_3.

11) Observation**:** Colour of the solution turns purple or blue.

Explanation:

- The (concentration) of Cl^- is increased.
- Equilibrium shifts to decrease concentration of Cl^-.
- Eqb shifts in favour of reactant.

More blue $[CoCl_4]^{2-}$ formed.

$[CoCl_4]^{2-}$

• equilibrium shifts to reduce temperature increase
• ΔH –ve/reaction exothermic
• reverse direction endothermic
• equilibrium shifts in endothermic/reverse direction.

12) Lighter brown / brown colour becomes less intense.

When the mixture is heated the endothermic reaction /absorption of heat is favoured. This is reverse reaction.

So amount of brown NO_2 gas is decreased / N_2O_5 increased so that the observed colour gets lighter

Lighter brown / brown colour becomes less intense.

As the pressure is increased the formation of fewer moles of gas is favoured.

This favours the reverse reaction since there are 5 moles of product gas compared with 2 moles of reactant gas. Thus the amount of brown NO_2 gas is decreased / N_2O_5 increased, so that the observed colour gets lighter.

13) $\dfrac{[FeSCN^{2+}]}{[Fe^{3+}][SCN^-]}$

Colour lightens / disappears / goes paler / more orange.

Removal of Fe^{3+} causes equilibrium position to shift towards the reactants in order to minimise the change, by replacing some of the Fe^{3+} that has been removed.

The new equilibrium mixture will therefore have less $FeSCN^{2+}$ and will be lighter in colour.

$\dfrac{[NH_3]^2}{[N_2][H_2]^3}$

Increased pressure of the system causes a shift to the right in order to decrease the pressure by forming fewer moles of gas. Therefore, the amount of NH_3 increases.

As the temperature is increased the amount of NH_3 produced decreases, indicating a shift to the reactants. As increasing temperature causes equilibria to shift in the endothermic direction, the forward direction (the reaction producing NH_3) must be exothermic.